**Solutions to Self-Check Questions**

**Question 1**

**(a)** $Cu(OH)_2 \bullet nCuCO_3 \rightarrow (n+1)\ CuO + H_2O + nCO_2$ **[1]**

**(b)**

1. Using an analytical/weighing balance, weigh and record the mass of a clean, empty and dry crucible.
2. Weigh out accurately about 5.00 g of solid $Cu(OH)_2 \bullet nCuCO_3$ into the crucible. Record the total mass of the crucible and the solid $Cu(OH)_2 \bullet nCuCO_3$.
3. Using a Bunsen burner, heat the crucible and its contents gently at first, and then heat strongly for 10 minutes.
4. Cool and weigh the crucible and its contents and record the total mass.
5. Repeat the heating-cooling-weighing process until constant mass is obtained.

**[1]** weigh using an analytical balance and record mass of empty crucible, and total mass of crucible with solid

**[1]** heat crucible and contents gently then strongly using a Bunsen burner

**[1]** allow crucible and contents to cool down before weighing and recording the mass

**[1]** repeat the heating-cooling-weighing process until constant mass is obtained.

Tabulation of results

| | |
|---|---|
| Mass of empty crucible / g | **A** |
| Mass of crucible and $Cu(OH)_2 \bullet nCuCO_3$ / g | **B** |
| Mass of crucible and its contents after first heating / g | **C** |
| Mass of crucible and its contents after second heating / g | **D** |
| Mass of crucible and its contents after third heating / g | **D** |

**[1]** Table must be constructed with correct headers and units.

Amount of CuO obtained = $\frac{\mathbf{D-A}}{63.5 + 16.0} = \frac{\mathbf{D-A}}{79.5}$ mol **[1]**

Amount of Cu in $Cu(OH)_2 \bullet nCuCO_3$ = amount of CuO obtained

% by mass of Cu = $\frac{63.5 \times \mathbf{(D-A)}}{79.5 \times \mathbf{(B-A)}} \times 100\%$ **[1]**

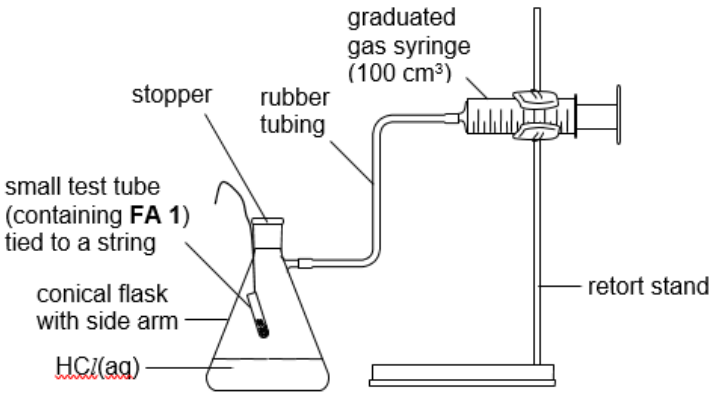
**Question 2**

**(a)** $\frac{\pm 0.05 \times 2}{23.50} \times 100\% = \pm 0.426\%$ **[1]**

**(b)** $CO_2(g)$ will dissolve / is soluble in water, resulting in a lower volume of $CO_2(g)$ measured. **[1]**

**(c)** The percentage purity calculated will be lower than expected. **[1]** The smaller the volume of $CO_2(g)$, the smaller the amount of $CO_2$ calculated, which in turn gives a smaller amount of $Na_2CO_3$ calculated in **FA 1** and hence lower percentage purity. **[1]**

**(d)(i)**




**[1]** Collection of gas using a gas syringe
**[1]** **FA 1** in small test-tube in conical flask

**(d)(ii)** With a gas syringe of 100 $cm^3$ capacity, assume that 75.0 $cm^3$ of $CO_2$ gas is collected.

Assume $V_{gas}$ to be 50 – 90% of gas collection apparatus capacity.

Amount of $CO_2$ gas = 75.0 / 24000 = $3.125 \times 10^{-3}$ mol **[1]**

$Na_2CO_3(s) + 2HCl(aq) \longrightarrow CO_2(g) + 2NaCl(aq) + H_2O(l)$

Amount of $Na_2CO_3 = 3.125 \times 10^{-3}$ mol
Molar mass of $Na_2CO_3$ = 2(23.0) + 12.0 + 3(16.0)
= 106.0 g $mol^{-1}$
Mass of $Na_2CO_3 = (3.125 \times 10^{-3})(106.0) = 0.3313$ g

Assuming that **FA 1** contains 90% by mass of $Na_2CO_3$,
mass of **FA 1** to be used = 0.3313 / 90 x 100 = 0.368 g **[1]**

## Question 3

**(a)** Unlike HC*l* which is a strong acid, $CH_3COOH$ is a weak acid and does not dissociate completely in aqueous solution. As energy is required to further ionise the unionised $CH_3COOH$, this resulted in a less exothermic enthalpy change of neutralisation.

**(b)** Pre-calculations
Let volume of **FA 1** or **FA 2** used be 40 $cm^3$.
n(HC*l*) = 40/1000 x 0.50 = 0.0200 mol
$n(CH_3COOH)$ = 40/1000 x 1.00 = 0.0400 mol

$n(Ba(OH)_2)$ = ½ x n(HC*l*) = 0.0100 mol
Volume of $(Ba(OH)_2)$ = 0.0100 / 0.5 x 1000 = 20 $cm^3$
(Note: Volume of $Ba(OH)_2$ should not exceed 20 $cm^3$ so that $Ba(OH)_2$ will be the limiting reagent.)

*(Assuming capacity of Styrofoam cup is 200 $cm^3$, volume of solution should be sufficient to submerge the bulb of the thermometer and should preferably not exceed more than half the capacity of the Styrofoam cup (i.e. up to 100 $cm^3$))*

Procedure

1. Using a 50 $cm^3$ measuring cylinder, add 20 $cm^3$ of $Ba(OH)_2$ into a clean and dry Styrofoam cup supported in a 250 $cm^3$ beaker.
2. Using a thermometer, measure and record the steady initial temperature of $Ba(OH)_2$.
3. Using another 50 $cm^3$ measuring cylinder, measure 40 $cm^3$ of **FA 1**. Ensure that both **FA 1** and $Ba(OH)_2$ are at the same initial temperature.
   *(Note: A burette/pipette cannot be used for measurement of FA1 as FA1 must be added to $Ba(OH)_2$ 'at one go'. Time is required for solution to run down the burette/pipette and is hence not suitable for use here)*
4. Add the **FA 1** into the cup containing $Ba(OH)_2$. Stir gently with the thermometer and record the highest temperature reached.
5. Repeat steps 1 to 4 by replacing **FA 1** with **FA 2**.
6. The acid which gave a lower temperature rise will be $CH_3COOH$ while the acid which gave a higher temperature rise will be HC*l*.

Key points:
- ✓ Use of appropriate apparatus & capacity (e.g. 50 $cm^3$ measuring cylinders) to measure the volumes of solutions, thermometer to measure temperatures, Styrofoam cup supported in 250 $cm^3$ beaker.
- ✓ Record steady initial temperature of $Ba(OH)_2$ & ensuring both solutions are at the same initial temperature.
- ✓ Record highest temperature reached.
- ✓ Determine identity of solutions based on temperature measurements.

**(c)** $H_2SO_4$ is a dibasic acid and HC*l* is a monobasic acid. Compared to HC*l*, per mole of $H_2SO_4$ reacts with NaOH to form 1 more mole of $H_2O$, releasing greater heat of neutralisation. Hence the enthalpy change of reaction per mole of $H_2SO_4$ is more exothermic than that of HC*l*.

**Question 4**

**(a)** $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$ **[1]**

**(b)** Pre-calculations **[1]**

Assuming that 20.00 $cm^3$ of $Ba(OH)_2$ was used,

Amount of $OH^- = \frac{20}{1000} \times 1.00 \times 2 = 0.04$ mol

Note:
Since extrapolation of a straight line is required, aim for ≥ 5 points before and after the estimated end-point to allow for good extrapolation.

The use of **5 $cm^3$ portions will allow 5 points before the end-point** and ending at 50.00 $cm^3$ will **allow 5 points after the end-point**.

Amount of $HNO_3$ used = 0.04 mol

Assuming $[HNO_3]$ = 1.5 mol $dm^{-3}$, volume of $HNO_3$ required = $\frac{0.04}{1.5} \times 1000 = 26.67$ $cm^3$

Thus, $HNO_3$ will be added in 5.00 $cm^3$ portions until 50.00 $cm^3$ has been added.

Procedure **[3]**

1. Using a 50.00 $cm^3$ burette, transfer 20.00 $cm^3$ of $Ba(OH)_2$ into a polystyrene cup supported in a 250 $cm^3$ beaker.
2. Place a thermometer into the polystyrene cup and measure the initial steady temperature of the $Ba(OH)_2$ solution in the cup.
3. Fill a separate 50.00 $cm^3$ burette with $HNO_3$ and place the polystyrene cup with the beaker under the burette.
4. Add 5.00 $cm^3$ of $HNO_3$ from the burette into the polystyrene cup. Stir the mixture with the thermometer. Measure and record the maximum temperature reached, $T_{max}$, attained and record it in the following table.

| volume of $HNO_3$ added / $cm^3$ | $T_{max}$ / °C |
|---|---|
| 0.00 | |
| 5.00 | |
| 10.00 | |
| 15.00 | |
| 20.00 | |
| 25.00 | |
| 30.00 | |
| 35.00 | |
| 40.00 | |
| 45.00 | |
| 50.00 | |

5. Repeat step 4 until a total of 50.00 $cm^3$ of $HNO_3$ has been added.

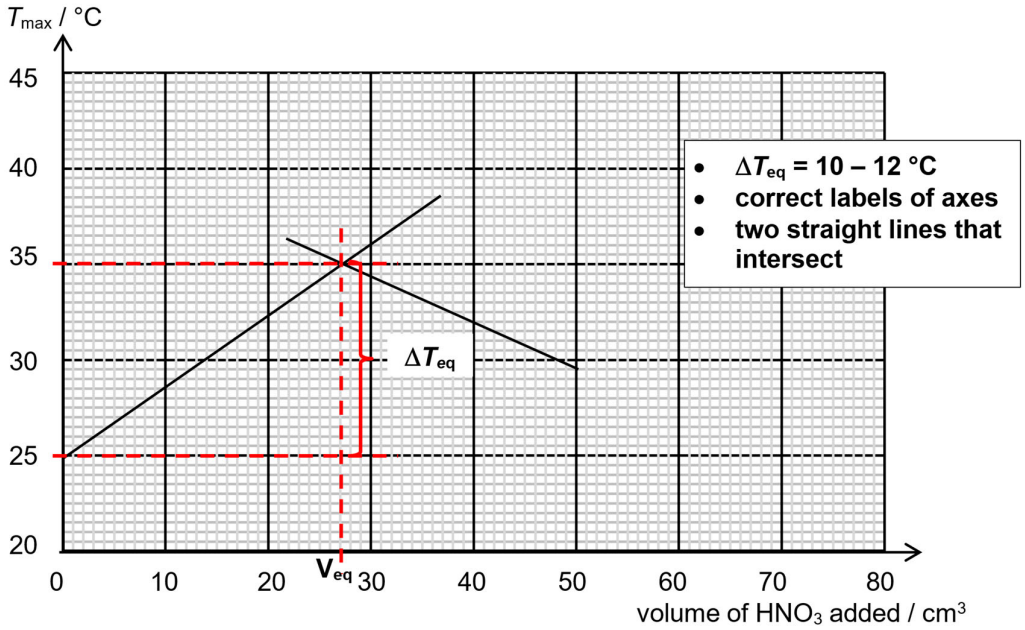
Key points:
- ✓ Burette to add $HNO_3$ + burette for measurement of $Ba(OH)_2$
- ✓ Styrofoam/polystyrene cup (for insulation)
- ✓ Thermometer + measurement of initial tempetaure & max temperature upon each addition of $HNO_3$
- ✓ Tabulation of data

How you would recognise that the equivalence-point has been passed **[1]**

Neutralisation is an exothermic reaction and the maximum temperature is reached at the end-point. The temperature of the mixture increases until the end-point is reached. After the end-point, the temperature will decrease as excess acid is added, because no further reaction is taking place and the excess $HNO_3$ added is at a lower temperature than the reaction mixture.

Sketch of the expected graph **[1]**

(Plot a graph of $T_{max}$ against volume of $HNO_3$. Draw best-fit lines for the points before and after the maximum $T_{max}$ has been reached and extend them until they intersect. The point at which the two best-fit lines meet corresponds to the volume of $HNO_3$ required for neutralisation and the maximum $T_{max}$.)




(Alternatively, graph of $\Delta T$ against volume of $HNO_3$ can be drawn ⇒ *Refer to Y5 Expt 7*)

An explanation of the shape of your graph **[2]**

Before equivalence point has been passed: $Ba(OH)_2$ is in excess while $HNO_3$ is limiting. As more $HNO_3$ is added, more $Ba(OH)_2$ is reacted, more heat is given off and the temperature rises.

After equivalence point has been passed: $HNO_3$ is in excess and $Ba(OH)_2$ has been completely reacted. No further reaction takes place with the addition of more $HNO_3$, hence no additional heat is given off. The temperature will decrease as excess $HNO_3$ is added, because no further reaction is taking place and the excess $HNO_3$ is at a lower temperature than the reaction mixture.

Treatment of data to find $[HNO_3]$ and value of $\Delta H_n$

Let $V_{eq}$ $cm^3$ be the volume of $HNO_3$ required to reach end-point and $\Delta T_{eq}$ °C be the maximum temperature rise obtained.

Amount of $OH^-$ = 0.04 mol (from pre-calculations)
Amount of $HNO_3$ to neutralise $Ba(OH)_2$ = 0.04 mol

$$[HNO_3] = \frac{0.04}{V_{eq}} \times 1000 = \frac{40}{V_{eq}} \text{ mol dm}^{-3} \textbf{ [1]}$$

Assume density of solution to be 1.00 g $dm^{-3}$ and specific heat capacity to be 4.18 J $K^{-1}$ $mol^{-1}$.

$$q = (V_{eq}+20)(1.00)(4.18)(\Delta T_{eq}) \text{ J}$$

$$\Delta H_n = -\frac{(V_{eq}+20)(1.00)(4.18)(\Delta T_{eq})}{0.04 \times 1000} \text{ kJ mol}^{-1} \textbf{ [1]}$$

**(c)** The maximum temperature rise will remain unchanged.

q = total mass of solution x specific heat capacity x temp change

When the volume of each solution used is doubled, the number of moles of water formed will double and the heat released will also be doubled. However, the total volume and hence mass of solution is doubled too. Therefore, the maximum temperature rise is the same. **[1]**

**OR**

$$\Delta H_n = -\frac{mc\,\Delta T}{n_{H_2O}}$$

$$\Delta T = \left(-\frac{\Delta H_n}{c}\right)\left(\frac{n_{H_2O}}{m}\right)$$

c is a constant and the value of $\Delta H_n$ for this reaction has a fixed value.

Hence, $\left(\frac{\Delta H_n}{c}\right)$ is a constant.

When the volume of each solution used is doubled,

- $n_{H2O}$ is doubled (as double the water is produced with double the reactants)
- total mass (m) of the solution will be doubled.

$\Rightarrow \frac{n_{H_2O}}{m}$ is also a constant.

Hence, $\Delta T$ remains unchanged.

**Question 5**

**(a)** Procedure **[5]**

1. Using a scalpel and an electronic balance, cut out and weigh five pieces of liver of the same size and mass.
2. Fill a burette with the $KMnO_4$ solution provided.
3. Using a 100 $cm^3$ measuring cylinder, add 100.0 $cm^3$ of the $H_2O_2$ solution provided to a 250 $cm^3$ conical flask labelled reaction mixture.
4. Place this conical flask into a thermostatically controlled water bath set at 50 $^oC$, with the $H_2O_2$ solution immersed in the water bath.
5. Place a thermometer into the same conical flask.
6. When the temperature of the $H_2O_2$ solution reaches 50 $^oC$, add one piece of liver from step 1 into the same conical flask. Start the stopwatch and swirl the mixture thoroughly to mix its contents while keeping the mixture still immersed in the water bath.
7. Using a 50 $cm^3$ measuring cylinder, add 50.0 $cm^3$ of the $H_2SO_4$ solution provided to a second 250 $cm^3$ conical flask.
8. Use a dropping pipette to transfer 10.0 $cm^3$ aliquot of the reaction mixture to a 10 $cm^3$ measuring cylinder.
9. Immediately transfer this aliquot into the second conical flask and vigorously swirl the mixture. Read and record the time of transfer in minutes and seconds, to the nearest second, when the aliquot is added.
10. Immediately titrate the $H_2O_2$ in the second conical flask against the $KMnO_4$ solution. The end-point is reached when the solution turns from colourless to a permanent pale pink. Record the titration results.
11. Wash out the second conical flask with water.
12. Repeat steps 7 to 11 four more times, at approximately 5-minute interval. The total reaction time should not exceed 25 minutes for this experiment.
13. Empty the conical flask labelled reaction mixture and wash it with water.
14. Repeat steps 2 to 13 four more times, each time with the temperature of the water bath in step 4 set at a different value of 10 $^oC$, 20 $^oC$, 30 $^oC$ and 40 $^oC$.

Key Points:
- ✓ Similar procedure to Y6 Expt 15 - appropriate apparatus and capacity; appropriate volumes of $H_2O_2$, $H_2SO_4$ and aliquot; starting stopwatch when liver (catalyst) is added; end-point colour change, recording of time and titration results etc
- ✓ Each piece of liver is of the same mass and size
- ✓ Incubating reaction mixture in a water bath with a thermometer
- ✓ Quenching at appropriate time intervals (at least 5 aliquots)
- ✓ Repeat procedure at different temperatures (at least 5)

Data manipulation **[2]**

15. For each experiment (at different temperature), plot a graph of the volume of $KMnO_4$ used against the decimal time, $t_d$, in minutes, to 0.1 min.
16. Extrapolate each graph to $t_d$ = 0.0 min
17. Draw a tangent to each graph at time $t_d$ = 0.0 min and find its gradient.
18. For each experiment (at different temperature),

initial rate = ½ (– gradient) $(0.0200 \times 10^{-3})$ $(\frac{5}{2})$ $(\frac{1}{10 \times 10^{-3}})$ *[Refer to Y6 Expt 15 calculations]*

Key Points:
- ✓ Plot separate extrapolated graphs of volume of $KMnO_4$ against decimal time, $t_d$, in minutes
- ✓ Draw tangent to graphs at $t_d$ = 0.0 min and obtain gradient to determine initial rate

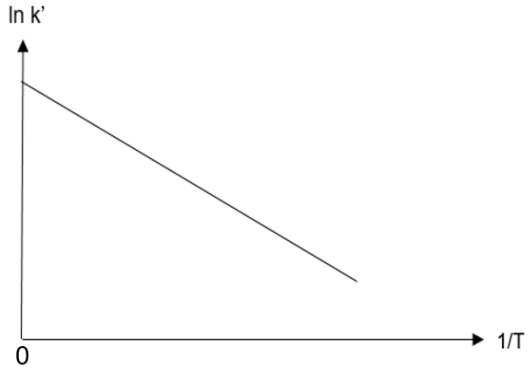
**(b)** For each temperature $T$ / °C,

- convert to $T$ / K and calculate $\frac{1}{T}$ / $K^{-1}$ **[1]**
- calculate the corresponding $k$' using the equation,
  $k' = \frac{\text{initial rate}}{0.170}$ where initial rate can be determined from each graph. **[1]**

Note:
initial rate = $k$' x initial $[H_2O_2]$
initial rate = $k$' x 0.170

$$k' = \frac{\text{initial rate}}{0.170}$$

**(c)**




**[1]**

Explanation:

$\ln k' = -\frac{E_a}{R}(\frac{1}{T}) + \ln A$

The graph fits the linear graph equation, y = mx + c.

Since both $E_a$ and R are positive, the graph is a downward sloping straight line with a negative gradient of $(-\frac{E_a}{R})$ and a y-intercept of ln A. **[1]**

**(d)** From the graph,

gradient = $-\frac{E_a}{R}$

$\therefore E_a$ = (gradient) (–R) = (gradient) (−8.31)

y-intercept = ln A

$\therefore$ A = $e^{(\text{y-intercept})}$ **[3]**